12-12-05 **Bonding** "Chpt 6 test commences here"

Types of Bonding [only care about an element's valence e-s]

• chemical bond - a link between atoms that results from mutual attraction of their nuclei for their e-s.

• Ionic bonds - forms as an electrostatic attraction between two oppositely charged ions.

ex. K· + :F: → K⁺ + :F:⁻ → KF [b/c ions attract one another]
(e⁻ transfer) •single arrow head → for a single e⁻.

(large difference in electronegativity)

→ most times → btw. metals & nonmetals
→ also between metals & polyatomic ion. → ex Ba^{2+} + SO_4^{2-} = $BaSO_4$ (both kind of bonding taking place.)

• covalent bonding - a link btw. atoms from e⁻ sharing.

ex H* + :Cl: → H:Cl: [Cl → slightly (-)]

ex 2 H· + H* → H:H [equal sharing]

→ doesn't have to be equal sharing ⇒ polarity

(not as big as a diff. in e⁻-neg.)

• covalent bonds exist btw. two nonmetals -or- w/in a polyatomic ion. [ex. SO_4^{2-}] O-S-O (with O above and below S)

Polarity in Bonds / Bond character : determined by difference in e⁻-neg. [in book & clean. structure]

	Difference of e-neg	ex
• nonpolar covalent bonds ⇒ equal sharing of e⁻s	0.0 - 0.3	H H [2.1 - 2.1 = 0]
• polar covalent bond ⇒ unequal e⁻ sharing	0.3 - 1.7	Cl - H [3.0 - 2.1 = 0.9]
• Ionic bond	1.7 - 4.0	Na - Cl [3.0 - 0.9 = 2.1]

12-14-05

Molecule - a group of two or more atoms joined by covalent bonds that can exist independently.

- for a covalent compound → smallest unit is a [molecule]
- for an element → smallest unit is an [atom]
- polyatomic atoms → don't exist by themselves ⇒ formula unit. (smallest ratio)

• How do bonds form?

- nuclei attract e⁻ ⇒ bond by sharing e⁻s
- distance of bond (length) is dictated by lowest PE. [avg. distance]

Bond E - E required to break a chemical bond.

Orbitals

ex. H· ↑ 1s [orbital notation] as atoms get closer & closer → e⁻ clouds touch.

H· ↑ 1s so... 1 e⁻ flips to allow for bond to take place. → happens b/c of repulsion. (change orientation/change direct)

[↑ 1s / ↓ 1s] ⇒ H:H (simplest → diatomic)

ex 2. :O: ↑↓ 2s ↑↓ ↑ ↑ 2p —2 flip→ ↑↓ 2s ↑↓ ↑ ↑ 2p] 2 bonds formed.

:O: ↑↓ 2s ↑↓ ↑ ↑ 2p ↑↓ 2s ↑↓ ↓ ↓ 2p ⇒ O::O

→ all sharing to get the Octet

Octet Rule - compounds form by atoms gaining, losing, or sharing e⁻s so that each atom has an octet in its outermost shell.

12-15-05

Covalent Bonds & Lewis Structures

ex: · CH_3I

Steps 1. Count total # of val. e⁻s.

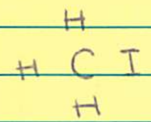
★ need to be able to find val. e⁻s fast!

ex. (cont)
$C \rightarrow 4\ e^{-}s \times 1 = 4e^{-}$
$H \rightarrow 1\ e^{-}s \times 3 = 3e^{-}$
$I \rightarrow 7\ e^{-}s \times 1 = 7e^{-}$
} total → $14e^{-}$
[hard # → has to be 14] (for this ex.)

2. Arrange elements into a "skeleton" structure [no e⁻s] w/ the least e-neg element @ center

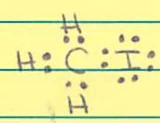
[hints: C is always central & H is never central]
(H: 1 val e⁻ ⇒ max of 2 e⁻ ⇒ can only make 1 bond.)

ex (cont 2)



3. ⓐ Between each atom place a pair of e⁻s for a bond.
ⓑ Around each nonmetal [not H] make sure there is an octet of e⁻s.

ex (cont 3)

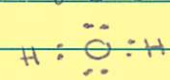


4. Check math.

ex (cont 4) started w/ 14 e⁻s --> 14 e⁻s to end ☺

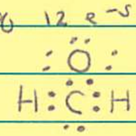
ex 2: H_2O step 1 → total of 8 e⁻s to "play w/"

step 2 →



ex 3: CH_2O step 1 → total of 12 e⁻s

step 2 →

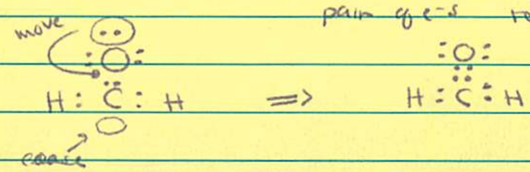


C → nonmetal
O → nonmetal

step 4 → started w/ 12 e⁻s → now have 14

5. If there are too many e⁻s
⇒ erase one pair of e⁻s on 1 atom & put another pair of e⁻s to make multiple bonds.

step 5 →



Rules
- H → always will have 1 bond & it's NEVER central
- C → always will have 4 bonds & no unshared e⁻ pairs & always (most x's) central
- N → always will have 3 bonds & 1 unshared pair of e⁻s
  unless it's positive then N can make 4 bonds. ex NH_4^+
- O → always will have 2 bonds & has 2 unshared pairs of e⁻s
  unless it's negative then O will make 1 bond & 3 unshared pairs of e⁻s.
  unless it's positive then O will make 3 bonds & 1 unshared pair of e⁻ → ex H_3O^+

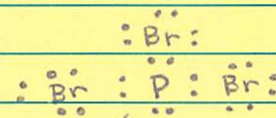
★ make it symmetrical when can.

test on 12-23-05 → up through 6-4.

12-19-05

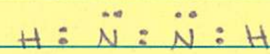
Lewis Structures Practice Worksheet

(1) PBr_3
↳ (5) + (7)(3)
= total # of 26 val. e⁻s

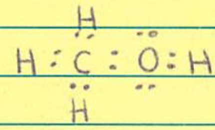


→ needs to fill octet.
[all nonmetals]
↳ same group as N (3 bonds w/ 1 pair of unshared e⁻s)

(2) N_2H_2
↳ (5)(2) + (1)(2)
= total # of 12 val e⁻s

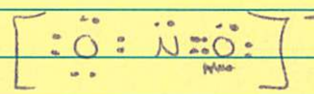


(3) CH_3OH [a structural formula - shows where structure is]



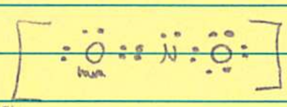
- molecular formula → CH_4O [no inclination of what structure looks like]

(4) NO_2^-
↳ ion → so add another e⁻

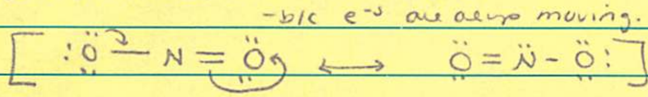


shows can't be by itself b/c it's an ion.

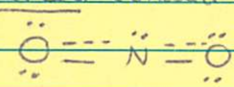
resonance = multiple structures possible for same molecule.
→ same as...



- both structures are correct & exist
- b/c e⁻s are always moving.

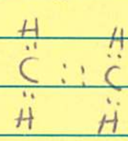


→ contributing structures (both contribute to overall structure)
- e⁻s are delocalized (shared btw. multiple atoms) [ex. of a fan running]



- $C_6H_{12}O_6$ → "glucose" - or fructose, etc.

(5) C_2H_4



12-20-05

Isomers - ~~multiple~~ same molecular formula but different structures.

12-21-05

Ionic Bonding → metals & nonmetals -or- metals & polyatomic ions

→ e^- transfer btw. e^{-s}

ex. Na [Ne] 1̲ (3s)

b/c Cl is more e^- neg ⇒ e^- transfers to Cl from Na

Cl [Ne] ⇅ (3s) ⇅ ⇅ ↿ (3p) ⟹ Na : [Ne] __ (3s) → Na^+

Cl : [Ne] ⇅ (3s) ⇅⇅⇅ (3p) → Cl^- chloride

• monoatomic ending → ide

• Na^+ - Cl^-

• charge goes off in all directions

• ions arrange in crystal lattice. (highly organized) repeating pattern

• ionic bonds → very strong, way more than covalent bonds.

• smallest covalent compound → molecule.

• smallest unit for ionic comp. → formula unit (shows smallest whole # ratio btw. ions)

ex. NaCl

$CuSO_4$

$CuCl_2$

bond E - amount of E to break covalent bonds.

lattice E - E required to disrupt crystal lattice

→ very high b/c more bonds in ionic compounds than covalent.

- El over charges.

Properties of Ionic Compounds

- high melting pt - high boiling pt. - conducts electricity → when dissolved in H_2O. (all ions)

- many free ions.

ex NaCl MP 801°C

$C_{12}H_{22}O_{11}$ 186°C

- very hard. - brittle (all b/c of very organized lattice)

- dissolve in H_2O easily (b/c H_2O is polar & attracted to charge of ions)

Metallic Bonding (btw. metals ...duh!)

↳ mix of ionic & covalent.

→ e^{-s} are delocalized (shared btw. multiple atoms) "e^- communism!"

- no one e^- belongs to one atom.

- shared ~~b/c~~ btw. the vacant d orbitals

⟹ "sea of e^{-s}"

- ductile
- malleable
- luster → b/c of e^- delocal
- tensile strength → b/c allowed to move.
- conductor of e^-
- current → flow of e^{-s}.

- covalent → b/c e^{-s} are shared btw. d orbitals
- ionic → b/c some e^{-s} are around nearby ~~to~~ make transitions.

01-04-06 — <u>2nd Part of Chpt 6</u>

<u>VSEPR</u> → Valence Shell e^- Pair Repulsion ⇒ outer shell e-'s repel one another.

↳ a theory used to predict 3-D shapes of molecules (therefore covalent bonds)

⇒ Bond ∠'s will form at the lowest PE (AKA least amount of e^- pair repulsion)

<u>VSEPR</u> (table)

Molecule type [general!]	# of unshared e- pairs on central atom	Geometry Type	Bond ∠ [tentative]	Ex. Molecule	Lewis Struc.	3D
A_2 (HCl, Hg_2^{2+})	N/A	linear	180°	H_2	H:H → H–H	H–H
AB_2 (HCN, SiF_2) — A ↑ an atom; B_2 ↑ w/ 2 branches	0	linear	180°	CO_2	:Ö=C=Ö:	Ö=C=Ö
AB_3	0	trigonal planar	120°	CO_3^{2-} → but... BH_3 (easier)	H–B–H with H below B	H–B–H with H above B
AB_4	0	tetrahedral	109.5°	CH_4	H–C–H with H above and below C	H–C–H, one H wedged (coming towards you!)
AB_3E (E ↑ unshared pair of e-'s)	1	trigonal pyramidal.	107°	NH_3	H–N̈–H with H below N	N̈ with three H, one wedged
AB_2E_2	2	bent	105.5°	H_2O	H–Ö–H; :Ö–H with H below O	:Ö–H with H below O

2nd 1/2 of Chpt 6 + Lewis x-structures + polarity → test/quiz Wed → 11m.

+ Chpt. hw's due.

01-06-06

- symmetry of a molecule w/ polar bonds ⇒ makes the molecule nonpolar.

01-09-06 Intermolecular Forces (IM) ex. CO_2

AKA Van der Waal's forces.

↳ [a force btw. molecules]

↳ the attractive force btw. molecules ✱ not bonds → just weak forces of attraction. → "rubbernecking"

– they are largely a result of bond polarity.

- strongest IM force → dipole-dipole forces

↳ an attractive force btw. dipoles.

ex H – Cl → polar b/c its difference in e-neg.

or

δ^{+} H – Cl δ^{-} δ → delta = partial charge.

– dipole → has 2 equal but opposite charged parts.

- H – Cl ..... H – Cl

opposite charges attract one another.

- BP

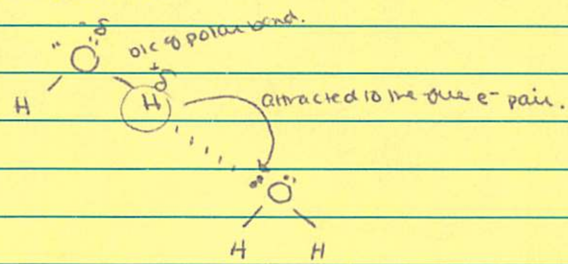
[nonpolar] Cl – Cl → – 34°C

[polar] I – Cl → 97°C ⇒ higher BP b/c of dipole attraction ⇒ more E required.

- Hydrogen bonds → ✱✱✱ THESE ARE NOT BONDS ✱✱✱

↳ a weak attractive force btw. a H bonded to a more e⁻-neg element & a free pair of e⁻s on another molecule.

ex.



- N_2 28 g/mol
- CO 28 g/mol
- CO_2 44 g/mol

} all gases.

- H_2O 18 g/mol → a liquid – b/c of H-bonds.

– water → makes life survive → Earth wouldn't exist → how it freezes. – less dense; → DNA & base bonds.

- London Dispersion forces – IM forces in nonpolar molecules.